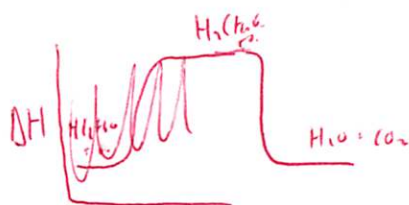
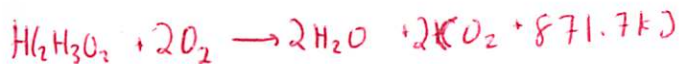


Thermochemical Equations

The complete combustion of pure, liquid acetic acid to form liquid water and carbon dioxide gas at constant pressure releases 871.7 kJ per mole of acid.

- Write a balanced thermochemical equation for this reaction.
- Draw an enthalpy diagram for the reaction.
- If 0.500 g of acetic acid is burned, what amount of energy is released?



$$0.500 \text{ g HOAc} \times \frac{1 \text{ mol HOAc}}{60 \text{ g HOAc}} \times \frac{-871.7 \text{ kJ}}{1 \text{ mol HOAc}} = \boxed{-7.26 \text{ kJ}}$$

Consider the following reaction, which occurs at room temperature and pressure:



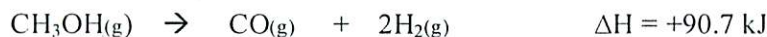
Which has a higher enthalpy under these conditions, 2Cl or Cl₂? Explain.

Since the Rxn is exothermic (-ΔH) 2Cl has a higher enthalpy.

$$\Delta H = H_{\text{Cl}_2} - H_{2\text{Cl}} = -243.4$$

↑ must be larger b/c neg ΔH.

Consider the following reaction:



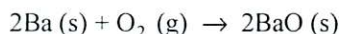
- Is heat absorbed or evolved in the course of this reaction? **absorbed.**
- Calculate the amount of heat transferred when 60.0 g of methanol is decomposed in the above reaction at constant pressure.
- For a given sample of methanol, the enthalpy change on reaction is 18.5 kJ. How many grams of hydrogen gas are produced?
- What is the value of ΔH for the reverse of this reaction? How many kilojoules of heat are released when 27.0 g of CO(g) reacts completely with H₂(g) to form CH₃OH(g) at constant pressure?

$$\text{ii } 60.0 \text{ g MeOH} \times \frac{1 \text{ mol MeOH}}{32 \text{ g MeOH}} \times +90.7 \frac{\text{kJ}}{\text{mol}} = \boxed{+170. \text{ kJ}}$$

$$\text{iii } 18.5 \text{ kJ} \times \frac{1 \text{ mol MeOH}}{90.7 \text{ kJ}} \times \frac{2 \text{ mol H}_2}{1 \text{ mol MeOH}} \times \frac{2.02 \text{ g H}_2}{1 \text{ mol H}_2} = \boxed{0.816 \text{ g H}_2}$$

$$\text{iv } = \boxed{-90.7 \text{ kJ}} \quad 27.0 \text{ g CO} \times \frac{1 \text{ mol CO}}{28.0 \text{ g CO}} \times \frac{1 \text{ mol MeOH}}{1 \text{ mol CO}} \times \frac{-90.7 \text{ kJ}}{1 \text{ mol MeOH}} = \boxed{-87.5 \text{ kJ}}$$

The value of ΔH° for the reaction below is -1107 kJ:



How many kJ of heat are released when 5.75 g of Ba reacts with 0.230 g of O₂ to form BaO(s)?

$$5.75 \text{ g Ba} \times \frac{1 \text{ mol Ba}}{137.3 \text{ g Ba}} = 0.420 \text{ mol Ba} \times \frac{1 \text{ mol O}_2}{2 \text{ mol Ba}} = 0.210 \text{ mol O}_2$$

$$0.230 \text{ g O}_2 \times \frac{1 \text{ mol O}_2}{32.0 \text{ g O}_2} = 0.00719 \text{ mol O}_2 \times \frac{-1107 \text{ kJ}}{1 \text{ mol O}_2} = \boxed{-7.96 \text{ kJ}}$$

Calorimetry

The molar heat capacity of a compound with the formula C_2H_6SO is $88.0 \text{ J/mol}\cdot\text{K}$. What is the specific heat of this substance?

$$\frac{88.0 \text{ J}}{\text{mol}\cdot\text{K}} \times \frac{1 \text{ mol } C_2H_6SO}{78.1 \text{ g } C_2H_6SO} = \boxed{1.13 \frac{\text{J}}{\text{g}\cdot\text{K}}}$$

Sp heat = $\frac{\text{J}}{\text{g}\cdot\text{K}}$

A sample of aluminum metal absorbs 9.86 J of heat, upon which the temperature of the sample increases from 23.2°C to 30.5°C . Since the specific heat capacity of aluminum is $0.90 \text{ J/g}\cdot\text{K}$, what is the mass of the sample?

$$q = mc\Delta T$$

$$m = \frac{q}{c\Delta T} = \frac{9.86 \text{ J}}{(0.90 \frac{\text{J}}{\text{g}\cdot\text{K}})(30.5^\circ\text{C} - 23.2^\circ\text{C})} = 1.5 \text{ g Al}$$

The specific heat of liquid bromine is $0.226 \text{ J/g}\cdot\text{K}$. How much heat (J) is required to raise the temperature of 10.0 mL of bromine from 25.00°C to 27.30°C ? The density of liquid bromine: 3.12 g/mL .

$$q = mc\Delta T$$

$$m = 10.0 \text{ mL} \times 3.12 \frac{\text{g}}{\text{mL}} = 31.2 \text{ g}$$

$$\Delta T = 27.30^\circ\text{C} - 25.00^\circ\text{C} = 2.30^\circ\text{C}$$

$$q = 31.2 \text{ g} \left(0.226 \frac{\text{J}}{\text{g}\cdot\text{K}} \right) (2.30^\circ\text{C}) = 16.2 \text{ J}$$

The ΔH for the solution process when solid sodium hydroxide dissolves in water is 44.4 kJ/mol . When a 13.9-g sample of NaOH dissolves in 250.0 g of water in a coffee-cup calorimeter, the temperature increases from 23.0°C to 37.8°C . Assume that the solution has the same specific heat as liquid water.

$$13.9 \text{ g NaOH} \times \frac{1 \text{ mol NaOH}}{40.0 \text{ g NaOH}} \times \frac{44.4 \text{ kJ}}{1 \text{ mol NaOH}} \times \frac{1000 \text{ J}}{1 \text{ kJ}} = 15400 \text{ J} = q$$

$$q = mc\Delta T$$

$$\Delta T = \frac{q}{mc} = \frac{15400 \text{ J}}{(250.0 \text{ g})(4.18 \frac{\text{J}}{\text{g}\cdot^\circ\text{C}})} = 14.8^\circ\text{C}$$

$$T_f = T_i + \Delta T = 23.0^\circ\text{C} + 14.8^\circ\text{C} = 37.8^\circ\text{C}$$

When 3.88 g sample of solid ammonium nitrate dissolves in 60.0 g of water in a coffee-cup calorimeter, the temperature drops from 23.0°C to 18.4°C . Calculate the ΔH (in kJ/mol NH_4NO_3).

$$q_{\text{soln}} = -q = mc\Delta T$$

$$= (60.0 \text{ g})(4.18 \frac{\text{J}}{\text{g}\cdot^\circ\text{C}})(18.4^\circ\text{C} - 23.0^\circ\text{C})$$

$$= -1153.4 \text{ J} = -1.15 \text{ kJ absorbed}$$

$$3.88 \text{ g NH}_4\text{NO}_3 \times \frac{1 \text{ mol NH}_4\text{NO}_3}{80 \text{ g NH}_4\text{NO}_3} = 0.0485 \text{ mol}$$

$$\Delta H = \frac{-1.15 \text{ kJ}}{0.0485 \text{ mol}} = -23.7 \frac{\text{kJ}}{\text{mol}}$$

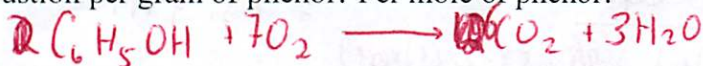
The British thermal unit (Btu) is commonly used in engineering applications. A Btu is defined as the amount of heat required to raise the temperature of 1.00 lb of water by 1.00°F. How many joules are in one Btu?

$$1.00 \text{ lb} \times \frac{454 \text{ g}}{1 \text{ lb}} = 454 \text{ g H}_2\text{O}$$

$$C = \frac{F - 32}{1.8} = \frac{1.8}{1.8}$$

$$q = \frac{4.18 \frac{\text{J}}{\text{g}^\circ\text{C}} (454 \text{ g}) (38^\circ\text{F} - 32^\circ\text{F})}{1.8} = 1052 \text{ J}$$

A 1.800 g sample of phenol ($\text{C}_6\text{H}_5\text{OH}$) was burned in a bomb calorimeter whose total heat capacity is 11.66 kJ/°C. The temperature of the calorimeter increases from 23.44°C to 30.57°C. Write a balanced chemical equation for the bomb calorimeter reaction. What is the heat of combustion per gram of phenol? Per mole of phenol?



$$q = 11.66 \frac{\text{kJ}}{^\circ\text{C}} \times (30.57^\circ\text{C} - 23.44^\circ\text{C}) = 83.13 \text{ kJ}$$

$$\frac{83.13 \text{ kJ}}{1.800 \text{ g}} = 46.19 \frac{\text{kJ}}{\text{g}}$$

$$\frac{46.19 \text{ kJ}}{\text{g}} \times \frac{94.0 \text{ g}}{1 \text{ mol}} = 4342 \frac{\text{kJ}}{\text{mol}}$$

A 50.0-g sample of liquid water at 25.0°C is mixed with 29.0 g of water at 45.0°C. What is the final temperature of the water?

$$q_{50} = -q_{29}$$

$$m \Delta T = -m \Delta T$$

$$m_{50} (T_f - 25^\circ\text{C}) = -m_{29} (T_f - 45^\circ\text{C})$$

$$50.0 \text{ g} (T_f - 25^\circ\text{C}) = -29.0 \text{ g} (T_f - 45^\circ\text{C})$$

$$50.0 T_f - 1250 = -29.0 T_f + 1305$$

$$79 T_f = 2555$$

$$T_f = 32.3^\circ\text{C}$$

Under constant-volume conditions the heat of combustion of benzoic acid ($\text{HC}_7\text{H}_5\text{O}_2$) is 26.38 kJ/g. A 1.540 g sample of benzoic acid is burned in a bomb calorimeter. The temperature of the calorimeter increases from 22.25°C to 27.20°C.

- What is the total heat capacity of the calorimeter?
- A 1.320 g sample of a new organic substance is combusted in the same calorimeter. The temperature of the calorimeter increases from 22.14°C to 26.82°C. What is the heat of combustion per gram of the new substance?
- Suppose that in changing samples, a portion of the water in the calorimeter was lost. In what way, if any, would this change the heat capacity of the calorimeter?

$$a \quad C = \frac{1.540 \text{ g} \times 26.38 \frac{\text{kJ}}{\text{g}}}{(27.20^\circ\text{C} - 22.25^\circ\text{C})} = 8.21 \frac{\text{kJ}}{^\circ\text{C}}$$

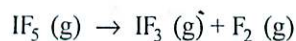
$$b \quad q_{\text{rxn}} = -C_{\text{cal}} \times \Delta T$$

$$= -8.21 \frac{\text{kJ}}{^\circ\text{C}} \times (26.82^\circ\text{C} - 22.14^\circ\text{C})$$

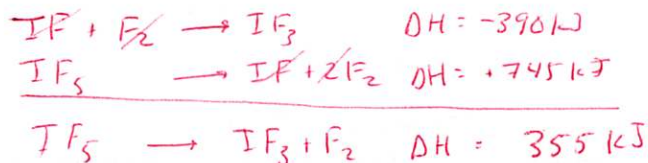
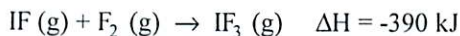
$$= -38.40 \text{ kJ}$$

$$\frac{-38.40 \text{ kJ}}{1.320 \text{ g}} = -29.00 \frac{\text{kJ}}{\text{g}}$$

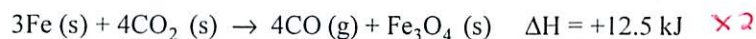
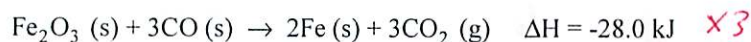
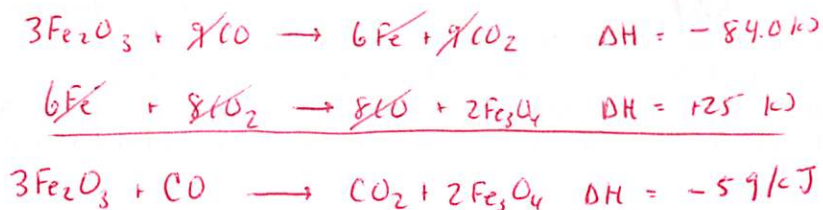
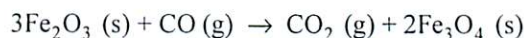
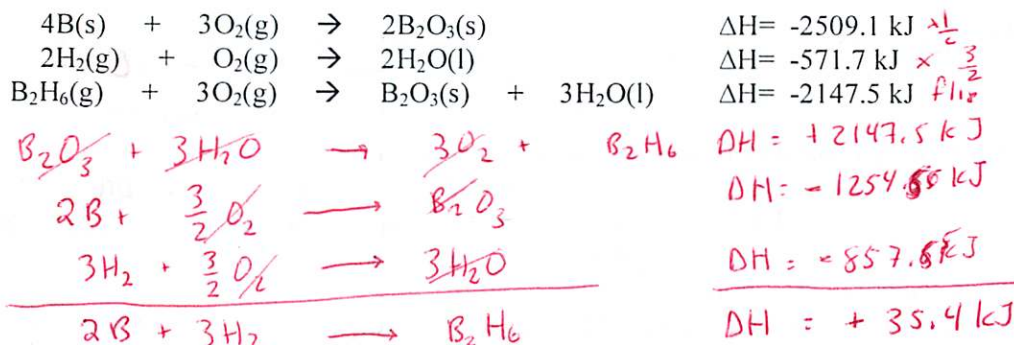
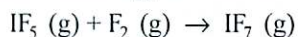
c The heat capacity would go $\uparrow \downarrow$
 B/c less H_2O means the T change would be greater so ΔT would higher in Eqn making C_{cal} lower.

What is the ΔH for the reaction below?

Use the following data to calculate your answer.



Given the following reactions:

calculate the enthalpy of the reaction of Fe_2O_3 with CO as shown below.Calculate the standard enthalpy of formation of gaseous diborane (B_2H_6) using the following thermochemical information.Given the data in the table below, calculate the $\Delta H^\circ_{\text{rxn}}$ for the following reaction.Substance $\Delta H_f^\circ (\text{kJ/mol})$

IF (g)	-95
IF ₅ (g)	-840
IF ₇ (g)	-941

$$\Delta H_{\text{rxn}} = \sum n(\Delta H_{\text{f, products}}) - \sum n(\Delta H_{\text{f, reactants}})$$

$$-941 \frac{\text{kJ}}{\text{mol}} - (-840 \frac{\text{kJ}}{\text{mol}} + 0) = -101 \frac{\text{kJ}}{\text{mol}}$$

The following reaction is known as the thermite reaction:



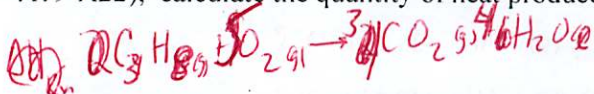
This highly exothermic reaction is used for welding massive units, such as propeller for large ships. Using enthalpies of formation values, Appendix 4 (pages A19-A22), calculate ΔH° for this reaction.

$$\Delta H_{\text{rxn}}^\circ = (-1676 \frac{\text{kJ}}{\text{mol}} + 2(0 \frac{\text{kJ}}{\text{mol}})) - (-826 \frac{\text{kJ}}{\text{mol}} + 2(0 \frac{\text{kJ}}{\text{mol}}))$$

$$= -850 \frac{\text{kJ}}{\text{mol}}$$

Propane

Many cigarette lighters contain liquid butane, $\text{C}_4\text{H}_{10}\text{(l)}$. Using enthalpies of formation, Appendix 4 (pages A19-A22), calculate the quantity of heat produced when 1.0 g of butane is completely combusted in air.



Propane

$$\Delta H_{\text{rxn}}^\circ = (4(-393.5 \frac{\text{kJ}}{\text{mol}}) + 5(-286 \frac{\text{kJ}}{\text{mol}})) - (104 \frac{\text{kJ}}{\text{mol}} + 5(0 \frac{\text{kJ}}{\text{mol}}))$$

$$= -2224 \frac{\text{kJ}}{\text{mol}}$$

$$1.0 \text{ g C}_4\text{H}_{10} \times \frac{1 \text{ mol C}_4\text{H}_{10}}{58.12 \text{ g C}_4\text{H}_{10}} \times \frac{-2224 \text{ kJ}}{1 \text{ mol C}_4\text{H}_{10}} = -38.3 \text{ kJ}$$

Complete combustion of 1 mol of acetone ($\text{C}_3\text{H}_6\text{O}$) liberates 1790 kJ:



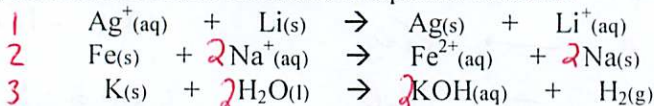
Using this information together with known heats of formation data, Appendix 4 (pages A19-A22), to calculate the heat of formation of acetone. If 5.0 grams of acetone is combusted, how much energy will be released?

$$-1790 \frac{\text{kJ}}{\text{mol}} = (3(-393.5 \frac{\text{kJ}}{\text{mol}}) + 3(-286 \frac{\text{kJ}}{\text{mol}})) - (X + 4(0 \frac{\text{kJ}}{\text{mol}}))$$

$$= 248.5 \frac{\text{kJ}}{\text{mol}} = X$$

$$5.0 \text{ g} \times \frac{1 \text{ mol C}_3\text{H}_6\text{O}}{58.0 \text{ g C}_3\text{H}_6\text{O}} \times \frac{-1790 \text{ kJ}}{1 \text{ mol C}_3\text{H}_6\text{O}} = -154 \text{ kJ}$$

Consider the following unbalanced redox reactions in aqueous solution:



$$\text{Li } \Delta H_f^\circ = -278.5 \text{ kJ}$$

$$\text{Fe}^{2+} \Delta H_f^\circ = -87.86 \text{ kJ}$$

- Balance each of the reactions.
- By using data in Appendix 4 (pages A19-A22), calculate the ΔH° for each of the reactions.
- Based on the ΔH° for the reactions, which do you expect to be unfavorable?
- Use an activity series to predict which of these reactions should occur. Are these results in accordance your conclusion in part c?

$$1 \Delta H = -278.5 \frac{\text{kJ}}{\text{mol}} - (+105 \frac{\text{kJ}}{\text{mol}}) = -383 \frac{\text{kJ}}{\text{mol}}$$

$$2 \Delta H = -87.86 \frac{\text{kJ}}{\text{mol}} - (2(40 \frac{\text{kJ}}{\text{mol}})) = -175.72 \frac{\text{kJ}}{\text{mol}}$$

$$3 \Delta H = 2(-481 \frac{\text{kJ}}{\text{mol}}) - (2(-286 \frac{\text{kJ}}{\text{mol}})) = -390 \frac{\text{kJ}}{\text{mol}}$$

d) Yes Li is higher than Ag so it will replace Ag. K will replace H₂O in liquid form.

c) B/c Rxns 1+3 have a (-) ΔH they will lead to products w/ a lower PE. so they are favorable.